Molar Mass Determination

An Introduction to Gas Laws

In this experiment, the molar mass of butane is determined using the ideal gas law and Dalton’s law of partial pressures.

Learning Objectives: Students will

- Learn how to use the ideal gas law and Dalton’s law of partial pressures to calculate the molar mass of a gas
- Practice collection of a gas using water displacement

Experiment Objectives:

- Calculate the molar mass of butane
- Statistically determine the accuracy and precision of the results

Introduction

Context: Texas Natural Gas

In 2009, Texas consumed 3.384 billion cubic feet of natural gas, more than any other state in the nation. Natural gas is a fossil fuel composed primarily of methane, although significant amounts of ethane, propane, and butane are present. Carbon dioxide, helium, hydrogen sulfide, and nitrogen gases are found in lesser amounts. During processing, the various gases are separated to form wet natural gas and dry natural gas. Wet natural gas is composed of hydrocarbons other than methane. Dry natural gas is composed of only methane. Household natural gas is dry natural gas. Wet natural gas is further refined, yielding ethane, propane, and butane gases. Butane gas is the subject of this lab.

<table>
<thead>
<tr>
<th>Component</th>
<th>Typical wt. %</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane (CH₄)</td>
<td>70–90</td>
</tr>
<tr>
<td>Ethane (C₂H₆)</td>
<td>5–15</td>
</tr>
<tr>
<td>Propane (C₃H₈) and Butane (C₄H₁₀)</td>
<td>&lt;5</td>
</tr>
<tr>
<td>CO₂, N₂, H₂S, etc.</td>
<td>balance</td>
</tr>
</tbody>
</table>
Molar Mass Determination

Chemical Principles: The Ideal Gas Law and Dalton’s Law of Partial Pressures

This experiment uses two gas laws to determine the molar mass of butane: the **ideal gas law** and **Dalton’s law of partial pressures**. Gas samples are described using four variables: Pressure (P), Volume (V), moles (n), and Temperature (T). The ideal gas law combines these 4 variables, allowing the calculation of one variable if the other three are known.

\[ PV = nRT \]

The proportionality constant, R, is the **universal gas law** constant. The numerical value of R depends on the units used for pressure and volume. Though there are various temperature scales, for gas law calculations, the temperature scale used is always the Kelvin temperature scale. When units for pressure are in atm and the volume units are in liters, the value of R is 0.08206 L atm/mol K. Measuring pressure and volume using units other than atm or L requires either the pressure and volume measurements to be converted to atm and liters, or the value of R is converted from atm and liters to the measurement units. For example, \( R = 62.34 \text{ torr L/mol K} \) when measuring pressure in torr and volume in liters. Similar conversions apply to volume measurements.

The second gas law is Dalton’s law of partial pressures. Dalton’s law states that for a gas mixture, the sum of the partial pressures for each gas in the mixture equals the overall pressure.

\[ P_{\text{total}} = P_A + P_B + \ldots \]

How does this apply to this experiment? During this experiment, butane gas released from a cigarette lighter will be collected by displacing water from a graduated cylinder. Before collection, the graduated cylinder is filled with water and inverted in a trough of water. As butane gas is bubbled into the graduated cylinder, the collected gas forces water out of the graduated cylinder (Figure 1).

![Gas collection over water.](image)

All liquids exert a vapor pressure above the bulk liquid. The vapor pressure increases as temperature increases. Water at 28°C has a higher vapor pressure than water at 20°C. **Table 1** lists the vapor pressure for water around room temperature. To correct for the pressure exerted by water, the vapor pressure of water must be subtracted from the total pressure of the gas mixture. The result is pressure of the collected gas, often referred to as the dry gas.

\[ P_{\text{dry}} = P_{\text{collected}} - P_{H_2O} \]
Table 1. Vapor pressure of water near room temperature.

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Vapor Pressure Water (torr)</th>
<th>Temperature (°C)</th>
<th>Vapor Pressure Water (torr)</th>
<th>Temperature (°C)</th>
<th>Vapor Pressure Water (torr)</th>
</tr>
</thead>
<tbody>
<tr>
<td>19.0</td>
<td>16.47</td>
<td>23.0</td>
<td>21.05</td>
<td>27.0</td>
<td>26.74</td>
</tr>
<tr>
<td>19.5</td>
<td>17.00</td>
<td>23.5</td>
<td>21.71</td>
<td>27.5</td>
<td>27.53</td>
</tr>
<tr>
<td>20.0</td>
<td>17.53</td>
<td>24.0</td>
<td>22.38</td>
<td>28.0</td>
<td>28.35</td>
</tr>
<tr>
<td>20.5</td>
<td>18.08</td>
<td>24.5</td>
<td>23.06</td>
<td>28.5</td>
<td>29.18</td>
</tr>
<tr>
<td>21.0</td>
<td>18.65</td>
<td>25.0</td>
<td>23.76</td>
<td>29.0</td>
<td>30.04</td>
</tr>
<tr>
<td>21.5</td>
<td>19.23</td>
<td>25.5</td>
<td>24.47</td>
<td>29.5</td>
<td>30.92</td>
</tr>
<tr>
<td>22.0</td>
<td>19.82</td>
<td>26.0</td>
<td>25.21</td>
<td>30.0</td>
<td>31.82</td>
</tr>
<tr>
<td>22.5</td>
<td>20.44</td>
<td>26.5</td>
<td>25.96</td>
<td>30.5</td>
<td>32.75</td>
</tr>
</tbody>
</table>

When the water level inside the graduated cylinder equals the water level in the trough, the gas pressure inside the cylinder equals the pressure outside the cylinder (the air pressure in the room). The pressure in the laboratory is read using a barometer. Your TA will read the barometer for you.

During this experiment, the butane is collected by downward displacement of water. As can be seen in the solubility table below, one of the reasons this technique works is because butane is essentially insoluble in water.

Table 2. Solubility of gases in water at 20°C.2

<table>
<thead>
<tr>
<th>Gas</th>
<th>Molecular Formula</th>
<th>Solubility*</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetylene</td>
<td>C₂H₂</td>
<td>0.117</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
<td>52.9</td>
</tr>
<tr>
<td>Butane</td>
<td>C₄H₁₀</td>
<td>0.061</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>CO₂</td>
<td>0.169</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>CO</td>
<td>0.0028</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>0.729</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>0.00016</td>
</tr>
<tr>
<td>Hydrogen sulfide</td>
<td>H₂S</td>
<td>0.385</td>
</tr>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td>0.0023</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>0.0019</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>0.0043</td>
</tr>
<tr>
<td>Sulfur dioxide</td>
<td>SO₂</td>
<td>11.28</td>
</tr>
</tbody>
</table>

* Grams of gas dissolved in 100 g of water, at 1 atm.

The reliability of a measurement can be assessed in one of two ways. The first considers the accuracy of the measurement, which describes how close the measurement comes to the true value. The reliability can also be analyzed by the measurement’s precision, which indicates how close a series of measurements are to each other. The accuracy of your determined molar mass will be assessed by computing the difference between your experimentally determined molar mass and the known molar mass of butane.
Molar Mass Determination

You will perform multiple trials of this experiment. This will enable you to verify that your results are reproducible and also determine the precision of your data by calculating the mean and deviation of your measurements. The mean, or average, denoted $\bar{x}$, is the sum of all values in a set divided by the number of values, $n$.

$$\bar{x} = \frac{x_1 + x_2 + x_3 + \ldots + x_n}{n}$$

The deviation of each measurement is the difference between the measurement, $x_i$, and the mean value $\bar{x}$.

$$\text{Deviation} = x_i - \bar{x}$$

More useful, though, is the standard deviation, $\sigma$. The standard deviation is a statistical measure of the amount of spread in a data series and can be calculated as shown below.

$$\sigma = \sqrt{\frac{(x_1 - \bar{x})^2 + (x_2 - \bar{x})^2 + \ldots + (x_n - \bar{x})^2}{n - 1}}$$

Focus

Through the application of the ideal gas law and Dalton’s law of partial pressures, students will learn how to calculate the molar mass of a gas. Analysis of data quality will enable the students to draw conclusions about the quality of their data.

Experiment

This is one of a handful of experiments in Chemistry 112 where a percentage of your grade will be based on the quality of your results.

A volume of butane gas will be collected by water displacement. The mass of butane in the gas sample will be determined from the loss in mass of the cigarette lighter. The pressure of the collected gas will be corrected for water vapor and the moles butane gas contained calculated. The molar mass of butane is then calculated from the mass and moles of butane.

Health, Safety, and Environmental (HS&E) Items

At all times, wear appropriate personal protection equipment (PPE), especially goggles for eye protection. Handle chemicals carefully. Butane from a cigarette lighter is used during this experiment and poses a fire hazard. Do not use matches, cigarette lighters, or strikers during this experiment. Perform this experiment under the benchtop fume hood.

Chemicals

Butane

Equipment

- a water trough
- 25 mL graduated cylinder
- top-loading balance
- small plastic tubing
- cigarette lighter
- barometer
- thermometer
Experimental Procedure

During this experiment, a total of four trials will be performed. The first trial you perform will be a practice run to refine your technique. The data from this practice run will be analyzed before proceeding through the rest of the experiment. This will give you the opportunity to make any necessary corrections to your technique. Subsequently, three trials will be performed where the precision and accuracy of your answer will dictate a portion of your grade.

Practice run

1. Record the barometric pressure and temperature of the room in your lab notebook.

2. Record, in your lab notebook, the mass of a butane cigarette lighter on an analytical balance to four digits past the decimal point. Attach one end of the plastic tubing to the gas outlet on the lighter.

3. Submerge a 25 mL graduated cylinder in the water trough. Without introducing air bubbles, invert the cylinder and slip the free end of the plastic tubing into the mouth of the graduated cylinder.

4. Depress the red lever on the cigarette lighter to allow butane gas to flow into the graduated cylinder. Collect 20–25 mL of gas. Set the gas pressure inside the graduated cylinder equal to the pressure in the room by moving the graduated cylinder up and down, without introducing air into the cylinder, so that the water levels inside and outside the graduated cylinder are equal. Record the volume of gas in the graduated cylinder in your lab notebook. Remember to estimate the volume to 1/10 of the smallest division marked.

5. Detach the plastic tubing connected to the butane lighter. Use an analytical balance to obtain the new mass of the lighter, to four digits past the decimal point, and record the mass in your lab notebook.

6. Calculate the molar mass of butane as described in steps 1–4 of the data reduction and analysis section below. Remember to account for the partial pressure of water vapor.

Experimental runs

1. Repeat steps 2 through 5 above. Perform three new trials of the experiment.

2. Return the lighters to the fume hoods and the graduated cylinders to the designated area on the bench in the back of the lab. Clean your workspace before leaving lab.

Post Lab Exercises

Data Reduction and Analysis

Data Reduction

1. Using the room pressure and the ambient temperature in the lab, look up the partial pressure of water vapor in the lab. Then, calculate the pressure of dry gas (butane) collected.

2. Use the ideal gas law to determine the number of moles of butane collected. Keep in mind that the units for pressure, volume and temperature need to match the units used for the universal gas constant, R, when performing this calculation.

3. Determine the mass of butane collected.
4. Using your answers from steps 2 and 3, calculate the molar mass of butane.

5. Repeat this process for each of the four trials (the initial test run plus the three real runs) of this experiment.

Statistical analysis

1. Calculate the average molar mass for your three experiment trials.

2. Calculate the percentage difference between your experimentally determined molar mass of butane and the real, or known molar mass of butane, 58.12 g/mol. The percent difference can be found using the equation shown below.

\[
\frac{|Known \; value \; – \; Experimental \; value|}{Known \; value} \times 100
\]

3. Determine the deviation of each trial from the average value.

4. Calculate the standard deviation of your data.

References


Data Reduction and Analysis

Experiment 7

Molar Mass Determination
An Introduction to Gas Laws

Name

Lab Partner

Section

TA Name

Score

☐ Data sheet stapled to back of DRA form.

Experiment parameters:

Room temperature __________ °C

Room temperature __________ K

Barometric pressure __________ torr

Partial pressure of water __________ torr

Calculate the pressure of dry gas collected.

General Equation:

Calculation:

Pressure of dry gas __________

Data, Practice Run:

<table>
<thead>
<tr>
<th>Practice Run</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial mass of lighter</td>
</tr>
<tr>
<td>Final mass of lighter</td>
</tr>
<tr>
<td>Mass butane used</td>
</tr>
<tr>
<td>Volume gas collected</td>
</tr>
</tbody>
</table>
Molar Mass Determination

Calculate the moles of butane collected.

**General Equation:**

<table>
<thead>
<tr>
<th>Calculation:</th>
</tr>
</thead>
</table>

Moles butane collected _____________

Calculate the molar mass of butane.

**General Equation:**

<table>
<thead>
<tr>
<th>Calculation:</th>
</tr>
</thead>
</table>

Molar mass butane ________________

Data, Experimental Runs:

<table>
<thead>
<tr>
<th></th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial mass of lighter</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final mass of lighter</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass butane used</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Volume gas collected</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Moles gas collected</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molar mass butane</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Difference (%) from actual molar mass</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Average molar mass from Trials 1–3 ________________
Calculate the percent difference between your experimentally determined molar mass of butane and the real molar mass.

**General Equation:**

Calculation:

Calculate the deviation of the molar mass for each trial.

**General Equation:**

Trial 1    |   Trial 2   |   Trial 3
---        |           |   |

Calculate the standard deviation of your data.

**General Equation:**

Calculation:

Standard Deviation ________________
Molar Mass Determination

Questions

1. Identify any sources of error in your experiment. Do not just say “human error;” cite specific examples based on the design or implementation of the experiment.

2. Could the molar mass of the following gases accurately be determined using the procedure in this experiment: ammonia (NH₃), carbon dioxide (CO₂), methane (CH₄) and nitrogen (N₂)? Explain why or why not.

3. Consider the scenario where 50.00 mg of copper and 10 mL of 0.2 M nitric acid were mixed. The reaction that occurs is: Cu(s) + HNO₃(aq) → Cu(NO₃)₂(aq) + NO₂(g) + H₂O(l). All of the gas formed from this reaction is collected over water in the same manner as the lab; the volume obtained was 20.52 mL. What percentage of the starting copper was consumed in this reaction? Assume room temperature (22.00°C) and atmospheric pressure (1.000 atm) for your calculations.